IGSI

UNIT - 8

CHEMISTRY FOR CLASS IX

CHEMICAL REACTIONS EQUILIBRIUM

राष्ट्रीय शैक्षिक अनुसंधान और प्रशिक्षण परिषद्
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PREFACE

The present series of twelve chemistry units has been developed for try-out of the Individually Guided System of Instruction (IGSI) in class IX. The description of IGSI and first two units of this series of units are available under a separate cover. This new system of instruction and the units have been developed along the lines of the National Policy on Education (NPE-86) and involve the participation of pupils in the process of learning. The units are suited for self-study with occasional help from a tutor. In the present context, these units will serve as examplar self-study material for secondary stage chemistry. In developing this unit, I was assisted by some of the chemistry teachers.

This unit contains an introduction for motivation, arousing interest, and to link the present unit with preceeding and next units. The objectives given in this unit are the expected learning outcomes, so that the pupil will know the ultimate goals he has to achieve. The suggested reading material provided in the unit guides the pupil to achieve pre-stated objectives. A number of intext and post-text questions, activities, and problems have been included to provide enough practice and chance for self evaluation.

The suggestions for the improvement of this unit will be welcomed.

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I. Introduction

In the last unit, we have studied that different chemical reactions proceed with different rates. Some of them like burning of fire-works proceed very fast. Others like rusting of iron proceed very slowly. Besides this, some of the reactions go to completion while others remain incomplete. The next question which comes to ones mind is how far a reaction can go ? Secondly, if a particular reaction does not go to completion, can we manage to carry it to completion ? If yes, then how ?

II Objectives

After completing this unit, you should be able to

(1) Explain reversible reactions.

(2) Explain the meaning of chemical equilibrium.

(3) Predict the effects of changing certain factors like temperature, pressure and concentration, on the state of equilibrium.

(4) Outline the significance of pH-scale and its applications.

(5) Identify acids and bases when their pH values are given.

III. Suggested Reading Material

You must have observed some children enjoying in the playground over a sea-saw. Two boys sitting on different sides balance each other and can attain a state of no change, This is a state of *equilibrium*. In chemical reactions also, such a state is attained when the concentrations of reactants and products remain unchanged. This state is known as chemical equilibrium. This equilibrium is however different from the equilibrium attained on sea-saw.

8.1 Reversible Reactions

Consider a familiar reaction between nitrogen and hydrogen to form ammonia,

$$3H_2 + N_2 \text{———} 2NH_3 + 94.3 \text{ KJ}$$

This is known as the forward reaction. Under the same conditions, ammonia decomposes to give back hydrogen and nitrogen.

$$2NH_3 \text{———} 3H_2 + N_2 - 94,3 \text{ KJ}$$

This is kmown as the reversible reaction.

Such a reaction which can take place in both forward and reverse directions is represonted as follows :

$$3H_2 + N_2 \underset{\text{reverse}}{\overset{\text{forward}}{\rightleftharpoons}} 2NH_3 + 94.3 \text{ KJ}$$

Such reactions in which both forward and reverse reactions occur are called reversible reactions. *All Chemical* reactions are infact reversible reactions. Many reactions, however, are said to go to completion because the reverse reaction does not take place to any noticeable extent. For example, when magnesium burns in oxygen, magnesium oxide is formed in 100% yield. No metallic magnesium is detectable because the magnesium oxide formed in the forward reaction does not appreciably decompose back into its elements For reactions that go nearly to completion, we use a single arrow ($\longrightarrow$) pointing towards the products. The double half head arrows sign ($\rightleftharpoons$) is used in between the reactants and products in representing reversible reactions.

$$2Mg + O_2 \longrightarrow 2MgO$$

8.2 Chemical Equilibrium

Let us consider the following reaction ;

$$A + B \rightleftharpoons C + D$$

As the reaction proceeds, concentration of A and B decreases whlie that of C and D increases. You have learnt in the previous unit that as the concentration decreases, the rate of the reaction decreases. So with the passage of time, rate of forward reaction decreases while the rate of reverse reaction increases. It can be represented graphically as shown in Figure 8 1.

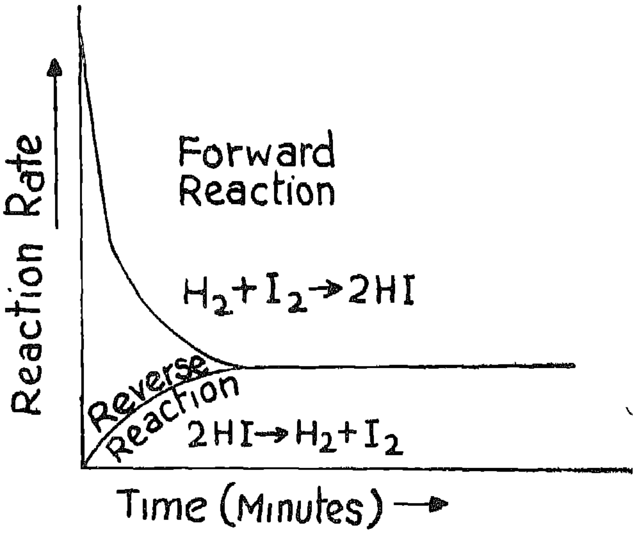

Fig. 8.1 Rates of formation and decomposition of hydrogen iodide.

What happens after a time t seconds ? After a time t seconds the rate of forward reaction becomes equal to the rate of reverse reaction. This stage is known as the stage of equilibrium. At equilibrium stage, as the forward and reverse reactions take place at equal rates, the composition of the equilibrium mixture (reactants and products) does not change with time. The equilibrium so attained is known as dynamic equilibrium The word dynamic here means that reaction continues to take place both in forward and reverse direction even after equilibrium is achieved But the equilibrium composition does not change In other words, at equilibrium the reactants keep on changing into products and the products into reactants, both at equal rates.

This dynamic Equilibrium state may be approached from either direction i. e. either from reactants side or products side. Consider the reaction of formation of ammonia.

This is represented as

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The equilibrium in the above reaction can be attained by starting the reaction either from N_2 and H_2 or from NH_3. The composition of all the three gases at equilibrium will be the same.

8 3 Dynamic Equilibrium in Physical Changes

The state of dynamic equilibrium also exists in physical changes such as dissolution of solids, vapourisation of liquids etc. When you add a small quantity of sugar to water at a certain temperature, sugar dissolves. This continues until the solution becomes saturated. If some more sugar is now added, it will remain undissolved. Here a dynamic equilibrium between dissolved and undissolved sugar molecule is established.

$$\text{Undissolved sugar (s)} \rightleftharpoons \text{dissolved sugar (aq)}$$

Similarly the dynamic equilibrium exists between liquid and vapours in a closed container at a constant temperature.

8.4. Ionic Equilibrium

Now let us see what happons when an ionic substance is dissolved in water. The strong electrostatic forces binding the ions become weak and the ions become free to move.

$$NaCl\,(s) + H_2O \longrightarrow Na^+\,(aq) + Cl^-\,(aq)$$

$$\text{Similarly } HCl\,(aq) \longrightarrow H^+\,(aq) + Cl^-$$

The dissociation of HCl, H_2SO_4 and HNO_3 into the respective ions is nearly complete. Therefore, these acids are known as strong acids. Weaker acids are dissociated to a lesser extent and a dynamic equilibrium exists between the dissociated and undissociated molecules of the acids For example

$$CH_3COOH \longrightarrow H^+ + CH_3COO^-$$

Similarly stronger electrolytes are dissociated into ions nearly to completion while weaker electrolytes dissociate to a lesser extent. The apparatus shown in Fig. 5.5, unit-5, can be used to distinguish between electrolytes and non-electrolytes and strong and weak electrolytes.

The solution of the compound to be tested is taken in a beaker The bulb glows only if the solution is an electrolyte. In case of strong electrolytes, the bulb glows brightly, and faintly with weak electrolytes. Substances like NaCl, KCl, $CuSO_4$ are strong electrolytes, NH_4OH, CH_3COOH are weak electrolytes. Urea and glucese are non-electrolytes.

8.5 Ionisation of Water

Pure water conducts electricity to a smaller extent. This means that pure water ionises into H^+ and OH^- ions to a very small extent. The undissociated water molecules are in equilibrium with its ions as depicted by the following equation.

$$\underset{\text{Water}}{H_2O(l)} \rightleftharpoons \underset{\text{Hydrogen ion}}{H^+ (aq)} + \underset{\text{Hydroxyl ion}}{OH^- (aq)}$$

The product of concentration of H^+ ions and OH^- ions released from water is a constant quantity at a constant temperature. It is equal to 10^{-14} mole^2 litre^{-2} at 25°C and is called the ionic product of water. Ionisation of pure woter produces equimolar concentrations of H^+ and OH^- ions.

Therefore,

$$[H^+] = [OH^-]$$

$$\text{Since } [H^+] \times [OH^-] = 1 \times 10^{-14} \text{ mol}^2 \text{ litrs}^{-2}$$

$$[H^+] = [OH^-] = \sqrt{10^{-14}} = 10^{-7} \text{ moles litre}^{-1}$$

On addition of an acid, $[H^+]$ increases. Therefore $[OH^-]$ should decrease since the product is always equal to 1×10^{-14} mole^2 litre^{-2} at 25°C. Similarly on addition of a base, $[OH^-]$ increases while $[H^+]$ decreases.

8.6 pH Scale

From section 8.5, we know that, in an acid solution, $[H^+] > [OH^-]$ or $[H^+] > 10^{-7}$ moles litre^{-1} and $[OH^-] < 10^{-7}$ moles litre^{-1}. In a basic solution, $[H^+] < [OH^-]$ or $[H^+] < 10^{-7}$ moles litre^{-1} and $[OH^-] > 10^{-7}$ moles litre^{-1} Thus we can say that

in neutral solution $[H^+] = 10^{-7}$ mol litre^{-1}

in acidic solution $[H^+] > 10^{-7}$ moles litre^{-1}

in basic solution $[H^+] < 10^{-7}$ mole litre^{-1}

H^+ ion concentration can vary from 1 to 10^{-14} mole litre^{-1} in common solutions. In 1919 a Swedish chemist S P. Sorensen realised that hydrogen ion concentration could be conveniently expressed by the numbers from 0 to 14 This is called the pH scale of the solution. pH is defined as the negative logarithm to the base 10 of the hydrogen ion concentration

$$\text{Mathematically } pH = \log \frac{1}{[H^+]} = -\log [H^+]$$

The set of positive numbers from O to 14 constitutes the pH scale.

For pure water $pH = -\log [H^+] = -(-7) = 7$

Acidic solutions have pH values less than 7

Basic solutions have pH values greater than 7

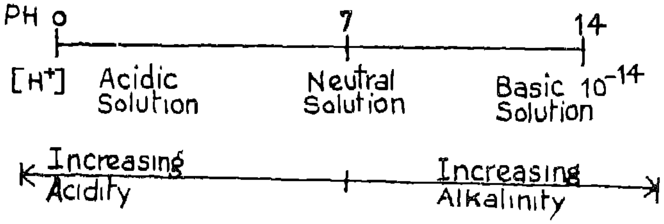

Fig. 8.9 pH Scale

Importance of pH

pH value finds important applications in agriculture. The pH of a soil controls the quality of the crop. Therefore pH of the soil has to be adjusted by the use of appropriate manures and fertilizers.

pH is also important in various biological processes The biological processes take place at specific pH. The [H^+] concentration of the processes like fermentation and enzyme hydrolysis is adjusted by determining their pH. The excess of acid in the stomach disturbs digestion. Various antacid solutions and tablets are used to decrease this excess acid.

pH value is also very important in chemistry. Many chemical reactions in analytical chemistry are carried out in a definite pH range.

Solved Examples of pH

1. Calculate pH of a solution of 0.01 M hydrogen ion concentration.

$$[H^+]=0.01\ M=\frac{1}{100}M=10^{-2}M$$

$$pH=-\log[H^+]=-\log 10^{-2}=2$$

2. Calculate pH of a solution whose $[OH^-]=10^{-10}$.

We know $[H^+][OH^-]=10^{-14}$

Substituting the value of $[OH^-]$ in the above equation,

$$[H^+]\times 10^{-10}=10^{-14}$$

$$pH=-\log(10^{-4})=4$$

3. Calculate $[H^+]$ of a solution whose pH = 5.

$$5=-\log[H^+]=\log\frac{1}{[H^+]}$$

$$\log\frac{1}{[H^+]}=5$$

$$\frac{1}{[H^+]}=\text{antilog } 5=10^5$$

$$[H^+]=\frac{1}{10^5}=10^{-5}M$$

8.7. Factors Affecting Equilibrium

Let us now study the effect of change of factors such as concentration, temperature, pressure and catalyst on the composition of the equilibrium mixture. We shall study this with the help of Lechatcher's principle

Lechatelier's Principle

The French Chemist Henry Louis LeChatelier (1850-1926) formulated a principle governing equilibrium. According to LeChatelier's principle, 'If a change is made in temperature, concentration or pressure of an equilibrium mixture, the equilibrium shifts in a direction so as to oppose the effect of that change'.

Let us now study the effects of above mentioned factors on equilibrium state one by one.

Effect of change in Pressure : The effect of change in pressure is observable in those cases where gases are involved in a reaction. Solids and liquids are not appreciably affected by changes in pressure. Consider an increase of pressure on reaction at equilibrium. According to LeChatelier's Principle, the equilibrium will shift in such a direction so as to oppose the effect of increase in pressure

Since pressure is directly proportional to the number of moles of gas or gases, the reaction will move in the direction in which the the number of moles of the gas will decrease. On decreasing the pressure the system will move in a direction in which the number of moles of the gas will increase. Let us take an example of dissociation of nitrogen tetraoxide.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

On decreasing the pressure, forward reaction will be favoured as it is accompanied by an increase in the number of moles of the gas. Therefore, yield of NO_2 gas would increase at low pressure

Consider the formation of ammonia.

$$N_2 + 3H_2 \rightleftharpoons 2\,NH_3$$

$$1+3=4 \text{ moles} \qquad 2 \text{ moles.}$$

The reaction is accompanied by decrease in the number of moles. Decrease in number of moles will decrease the pressure If the pressure on the reaction is increased, forward reaction will be favoured.

Effect of change in Temperature : Let us lstudy the conversion of NO_2 to N_2O_4 which is an exothermic reaction.

$$2NO_2(g) \rightleftharpoons N_2O_4(g) + 58.2 \text{ KJ}$$

On decreasing the temperature, the equilibrium will shift in a direction in which heat is evolved. In Other words the equilibrium of the reaction will be shifted towards the products side. This formation of N_2O_4 is favoured at low temperature. Similarly at higher temperature, endothermic reaction (reverse reaction) is favoured resulting in greater decomposition of N_2O_4 into NO_2.

Activity : To study the effect of change of temperature on a reversible chemical reaction.

$$2NO_2(g) \rightleftharpoons N_2O_4(g) + 58\ 2\ KJ$$

Procedure : Prepare NO_2 gas by adding copper turning to concentrated nitric acid. A brown coloured gas (NO_2) is obtained. Collect the gas in three test tubes. Cork the tubes and place them in a beaker containing water at room temperature for ten minutes. Now keep one of the test tubes in hot water (80°C) and the other in ice (0°C). Compare the colour intensity of the gas in the two test tubes with the third test tube kept in water at room temperature.

Questions

(1) At low temperature (O°C), what is the colour of the gas ?

(2) At high temperature (80°C), what is the effect on the intensity of the brown colour of gas ? Explain the observations.

(3) Consider the following reaction :

$$N_2 + 3H_2 \rightleftharpoons 2NH_3 + 94.3\ KJ$$

What will be the effect of decrease in temperature on the equilibrium state of the reaction ? Explain

Effect of Change in Concentration : If we increase the concentraion of a particular reactant or product of a reaction at equilibrium, the reaction will take place in the direction in which the added material gets consumed. Let us consider the following example :

$$\underset{\text{Ethyl acetate}}{CH_3\,COOC_2H_5(l)} + H_2O(l) \rightleftharpoons \underset{\text{Acetic acid}}{CH_3\,COOH(l)} + \underset{\text{Ethyl alcohol}}{C_2H_5OH}$$

If we increase the concentration of C_2H_5OH by adding some ethyl alcohol from outside, the reverse reaction takes place to as to decrease the concentration of added ethyl alcohol So the reverse reaction starts faster and the hydrolysis of ethyl acetate is decreases. Consider another reaction .

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

If ammonia gas is removed from the equilibrium mixture more and more of hydrogen and nitrogen will react to produce more of ammonia. This property is utilized in chemical industries to get a better yield of ammonia.

Effect of Catalyst : The catalyst does not affect the composition of the equilibrium mixture. The only advantage of using a catalyst in a reversible reaction is to reduce the time required to attain the equilibrium. By adding a catalyst to a reaction, equilibrium is attained quickly. The presence of a catalyst increases the rate of both forward and the reverse reactions equally.

Now let us try to apply the information/knowledge gathered from the Le Chatelier's principle to certain reversible processes taking place in chemical industries. One of them is the Haber's process for the manufacture of ammonia.

$$N_2 + 3H_2 \underset{\text{Endothermic}}{\overset{\text{Exothermic}}{\rightleftharpoons}} 2NH_2 + 94.3 \text{ K.J.}$$

We have to find out the temperature and pressure which gives the maximum yield of ammonia. This is a reversible exothermic reaction. High temperature would favour that reaction where heat energy is absorbed i.e endothermic reaction. So increase of temperature would favour decomposition of ammonia. Hence to get better yield of ammonia, reaction temperature should be lowered.

Low pressure favours that reaction where number moles of gases are increasing The total number of moles decreases in the formation of ammonia. There are 4 moles of reactants and 2 moles of products So low pressure would favour decomposition of ammonia. The yield of ammonia would increase if the pressure is high. Hence the optimum condition for the better yield of ammonia are high pressure and low temperature

Let us now consider the formation of sulphur trioxide in the *contact process* for the manufacture of sulphuric acid as another example.

$$2\ SO_2(g) + O_2\ (g) \rightleftharpoons 2\ SO_3 + \text{heat}$$

This is a reversible exothermic reaction. Lowering the temperature would favour the formation of sulphur trioxide High temperature favours endothermic reaction (reverse reaction).

In the formation of SO_3, we see from chemical equation that three moles of reactants produce two moles of products. Hence the reaction is accompanied by a net decrease in pressure. If the pressure on the reaction at equilibrium is increased, the reaction will take place in a direction so as to decrease the pressure. In other words the increase of pressure will favour forward reaction

IV. Home Assignment

Consider a mixture containing 1 mole of N_2 and 3 moles H_2 brought to equilibrium in presence of a catalyst at 777K and 10 atm. pressure. The mixture reacts to form ammonia according to the equation.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g + 94.3\ KJ$$

The yield of ammonia is very low. Only about 2.5% of the reactants are converted into prodicts. Suggest how can the yield of NH_3 be increased (a) by altering the pressure and (b) by altering the temperature.

V. Self Assessment

1. Fill in the Blanks.

(i) At equilibrium the rates of forward and reverse reactions are——————————.

(ii) In exothermic reactions at equilibrium, increase of temperature favours——————————reaction.

(iii) ————————of temperature of endothermic reaction at equilibrium favours forward direction.

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2. **Select true and false statements**

   (i) In reversible exothermic reactions the yield of the products can be increased by keeping the temperature low.

   (ii) In endothermic reactions the yield can be increased by keeping temperature low.

   (iii) Weak electrolytes are completely dissociated in aqueous solutions.

   (iv) For the reaction

   $H_2(g) + I_2(g) = 2HI(g)$

   at equilibrium, the increase of pressure will

   (a) increase the amount of HI.

   (b) decrease the amount of HI.

   (c) have no effect.

3 What is the relation between rates of forward and reverse reactions at equilibrium.

4. The pH value of two drinks A and B are 6 4 and 7.2 respectively. Which drink will be alkaline and which will be acidic.

5. Explain the following :

   (i) Chemical equilibrium.

   (ii) Le Chatelier's Principle,

   (iii) Reversible reaction.

   (iv) pH and pH scale

Teacher's Guide

The unit may be introduced by recalling a few examples of fast, moderate and slow reactions from unit - 7 A question may now be raised as to whether all reactions continue till all the reactants are completely consumed, keeping in view the possibility of the products combining to give back reactants. The forward and reverse reactions occurring simultaneously would result in an incomplete reaction. Reversible reactions may be explained with examples of suitable chemical reactions. Actually all reactions can occur in both the forward and reverse direction to some extent under certain reaction conditions. In case of the so called irreversible reactions, Keq lies largely in favour of products side.

The effect of changes of concentration, temperature and pressure on the state of equilibrium may be qualitatively explained with the help of Le Chatelier's principle.

In introducing pH, the use of logarithim table may be explained in the class if the students have not used the same earlier. The students may be asked to note the pH of a few fruits juices with the help of pH paper and draw the conclusion about the acidity of juices.